BTEC Assignment Brief

Qualification	Pearson BTEC International Level 3 Subsidiary Diploma in Applied Science Pearson BTEC International Level 3 Foundation Diploma in Applied Science Pearson BTEC International Level 3 Diploma in Applied Science Pearson BTEC International Level 3 Extended Diploma in Applied Science Pearson BTEC International Level 3 Extended Diploma in Biomedical Science Pearson BTEC International Level 3 Extended Diploma in Environmental Science
Unit number and title	**Unit 16: Applications of Inorganic Chemistry**
Learning aim(s)	**D:** Investigate reduction and oxidation reactions in order to understand analytical and industrial applications
Assignment title	**Reduction and oxidation reactions**
Assessor	
Issue date	
Hand in deadline	

Vocational Scenario or Context	You are a laboratory technician working for a chemical company that specializes in making inorganic compounds. In your role you need knowledge of how reduction and oxidation reactions (Redox reactions) play an important role in many industries and in our daily lives. The development of for example lead-acid, alkaline and lithium batteries and fuel cells which are used in transport and many other everyday applications. The development of batteries for electric cars, buses and lorries are crucial in the fight to reduce air pollution and the amount of carbon dioxide in the atmosphere to help stop global warming. A lot of electrical energy is now produced by solar cells during the daytime and research is taking place to improve batteries to increase the amount of electricity they can store which can be used at night. Your supervisor has set up some Continuing Professional Development (CPD) for you to ensure that your knowledge and skills remain up to date.

Task	As part of your annual CPD in your role of laboratory technician you are required to **investigate** concentrations of analytes and **compare** experimentally determined cell voltages with calculated values. Your supervisor has also asked you to prepare a report that **evaluates** the use of redox reactions and standard electrode potentials and **analyses** a range of redox reactions and electrochemical cells. Your final portfolio will contain the following 5 sections: Section 1 - **Compare** experimentally three selected electrochemical cells • Write out the half equations and redox equations for the following cells:

- $Zn_{(s)} \mid Zn^{2+}_{(aq)} \,\vdots\, Cu^{2+}_{(aq)} \mid Cu_{(s)}$
- $Zn_{(s)} \mid Zn^{2+}_{(aq)} \,\vdots\, Pb^{2+}_{(aq)} \mid Pb_{(s)}$
- $Pb_{(s)} \mid Pb^{2+}_{(aq)} \,\vdots\, Cu^{2+}_{(aq)} \mid Cu_{(s)}$

- Calculate the standard cell voltages for these cells using standard reduction potentials. Then set up and measure the actual voltage of the three cells (see "Measuring cell potentials" worksheet) and **compare** the measured voltages to the calculated voltages. Show your workings for the calculations and a clear, logical table for your measurements. In your conclusion to the experiment you must comment upon the accuracy of your practical results compared to your calculated voltages, accounting for any differences.
- **Analyse** a range of electrochemical cells: Plan four electrochemical cells from the following two half cells and their standard electrode potentials. Write down oxidation and reduction half equations and a labelled diagram of each of the four planned electrochemical cells:

  - $Fe^{2+} + 2e^- \rightarrow Fe$ Standard Potential E° (volts) = -0.41
  - $Cu^{2+} + 2e^- \rightarrow Cu$ Standard Potential E° (volts) = +0.34

  When planning your cells you will need to consider the possible types of electrochemical cells (electrolytic and galvanic) and the possible use of an inert electrode such as platinum.

Section 2 – **Analysis** of Redox equations

- Complete the "Oxidation/Reduction equation worksheet" and include this in your portfolio

Section 3 - **Investigate** the concentrations of analytes using redox procedures and colorimetry

- Using a standardised solution of $KMnO_4$, carry out the Fe^{2+}/MnO_4^- titration to determine the concentration of a solution of Fe^{2+} ions (from iron tablets or steel wool)

- Standardise a solution of sodium thiosulfate with potassium iodate, KIO_3. Using your standardised sodium thiosulfate solution, select two (or more) of the following titrations to carry out:
  - a **determination** of the concentration of iodine;
  - a **determination** of the concentration of Cu^{2+} ions (eg in brass);
  - a **determination** of the concentration of hydrogen peroxide;
  - a **determination** of the concentration of hypochlorite (ClO^-) ions (eg in bleach)

- For at least one of your analytes above, use colorimetry to determine the concentration of the analyte. Compare the accuracy of the concentration of the analyte determined by redox titration and by colorimetry.

  You will need to research and source appropriate methods, which you

	must risk assess and get approval from your assessor to carry out. Alternatively, your supervisor may choose to supply these for you. Write out the redox equations involved for each of the titrations and use these to help calculate the concentration of the analyte. Using oxidation numbers and numbers of electrons involved, **analyse** how this determines the stoichiometry of the equations and therefore how this has helped you determine the concentration of the analyte. Section 4 – **Evaluate** analytical techniques used in industry • Write a report **evaluating** the following analytical techniques used in industry: o Iodine/thiosulfate titration to determine the concentration of Cu^{2+}; o Iodine/thiosulfate titration to determine the peroxide value of fats and oils; o Iodine/thiosulfate titration to determine the ClO^- concentration in bleach; o Dichromate(VI) to determine chemical oxygen demand. Your evaluation should be based upon a variety of sources of information (including your own practical work). You will examine how accurate and reliable the techniques are, discuss how widely used the techniques are used in industry (specifying the type of industry) and suggest if there are other alternative analytical procedures that are, or could be used. Section 5 **– Evaluate** standard electrode potentials in commercial applications • **Evaluate** how standard electrode potentials are applied to a range of commercially available electrochemical cells (batteries and fuel cells) and methods of corrosion prevention (e.g. sacrificial protection) . Use your calculations and other factors such as cost of materials, durability and alternatives.
Checklist of evidence required	A report that includes: • An investigation of the concentration of analytes in solution and by colorimetry • A comparison of three measured cell voltages with calculated voltages using standard electrode potentials • An analysis of a range of redox reactions and electrochemical cells • An evaluation of the use and limitations of standard electrode potentials.

Criteria covered by this task:	
Unit/Criteria reference	To achieve the criteria, you must show that you are able to:
D.D4	Evaluate the use of redox reactions and standard electrode potentials in analytical and commercial applications

D.M7	Analyse a range of redox reactions and electrochemical cells in terms of oxidation numbers, equations and standard electrode potentials involved
D.P9	Compare experimentally determined cell voltages for metal/metal ion electrochemical cells with calculated values
D.P8	Investigate the concentration of analytes using redox titration and colorimetry
Sources of information to support you with this Assignment	All sources of information suggested are UK based. Centres are free to amend these sources to those available in the Country of delivery where more appropriate. https://www.chemguide.co.uk www.STEM.org.uk – Equilibrium III – Redox Reactions, Beer's Law Lab www.edu.rsc.org - Beer's law simulation, Redox equilibria starters Within the copyright laws learners could search YouTube for: electrochemical cells; introduction to redox reactions; colorimetry
Other assessment materials attached to this Assignment Brief	The following worksheets have been included as part of this assignment: 1. "Measuring cell potentials" worksheet 2. "Oxidation/Reduction equation worksheet"
Notes to the Assessor and technician	The practical activities suggested for this assignment are required to be safety checked by the Centre before use. Pearson Education can take no responsibility for the safety of any practical activities. Before doing any practical activity, you should always carry out your own risk assessment and refer to the necessary regulations. In particular, any local rules issued by your employer must be obeyed, regardless of what is suggested here. Centre should use Health and Safety information provided by a reputable source. In the UK such an example organization is CLEAPSS http://science.cleapss.org.uk/

Measuring cell potentials

Aim

The purpose of this experiment is to construct three electrochemical cells, and to measure the potential difference between the electrodes, noting the polarity.

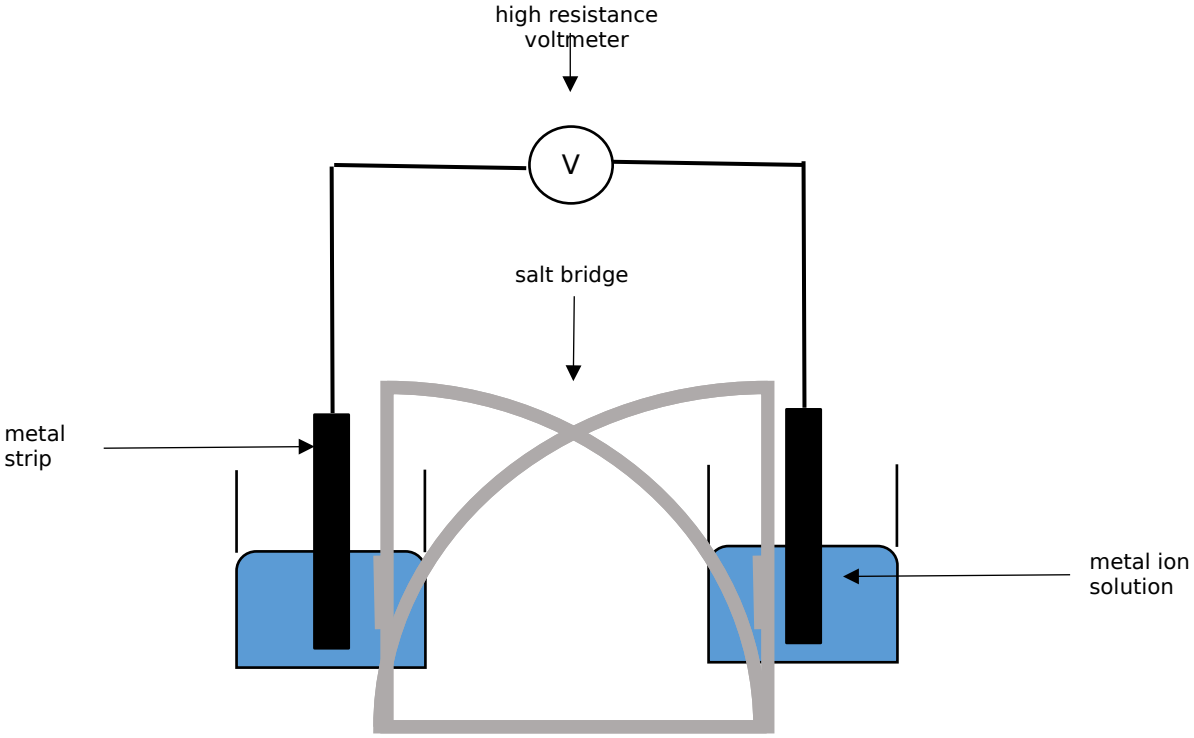
Introduction

You will construct the cells given below according to the method and measure the potential difference between the metal electrodes using a high resistance voltmeter.

- $Zn_{(s)} \mid Zn^{2+}_{(aq)} \,\vdots\, Cu^{2+}_{(aq)} \mid Cu_{(s)}$
- $Zn_{(s)} \mid Zn^{2+}_{(aq)} \,\vdots\, Pb^{2+}_{(aq)} \mid Pb_{(s)}$
- $Pb_{(s)} \mid Pb^{2+}_{(aq)} \,\vdots\, Cu^{2+}_{(aq)} \mid Cu_{(s)}$

When measuring the potential difference between the electrodes of each cell, the polarity of each electrode is indicated by the positive sign and negative sign on the voltmeter (or red and black terminals respectively). Thus, you can determine at which electrode electrons are needed and at which electrode electrons are produced. From this, you can work out the direction in which the overall reaction proceeds.

Diagram



Chemicals and Equipment

- copper(II) sulfate solution (0.1M)
- zinc sulfate solution (0.1M)
- lead(II) nitrate solution (0.1M)
- strip of copper foil (or wire)
- strip of zinc foil (or wire)

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- strip of lead foil (or wire)
- emery paper, 3 small pieces
- saturated potassium nitrate solution
- safety spectacles
- 4 beakers, 50 cm^3
- 3 measuring cylinders, 20 cm^3
- 3 connecting leads with crocodile clips attached
- filter paper strips
- voltmeter, high resistance

CAUTION:
Copper(II) sulfate solution is harmful.
Lead(II) nitrate is toxic.
You must wear safety spectacles throughout the experiment.

Method

1. If necessary, clean each metal strip (or wire) with a separate piece of emery paper.

2. Place each metal strip in a separate beaker. Hold each strip vertically against the inside of the beaker so that about 2 cm projects above the rim. Fold the projection down over the rim of the beaker and clamp it in position with a crocodile clip attached to a lead.

3. Pour approximately 20 cm^3 of the appropriate salt solution into each beaker and ensure that each metal strip dips into a solution of its own ions. Ensure that the crocodile clips remain dry.

4. Prepare a salt bridge by soaking a strip of filter paper in saturated potassium nitrate solution. Allow the surplus solution drain off by hanging the strip over the fourth beaker.

5. Connect the zinc and copper electrodes to a high resistance voltmeter, as shown in the diagram below and complete the circuit with the salt bridge.

6. If the reading on the voltmeter is negative, reverse the connections to obtain a positive reading. Record the potential difference in a suitably designed results table, noting which electrode is positive (connected to the red terminal of the voltmeter) and which is negative.

7. Remove the salt bridge as soon as possible and throw it away. Disconnect the voltmeter.

8. Repeat steps 4 to 7 for the other two cells.

9. Using appropriate standard reduction potential data, calculate a theoretical value for the potential difference of each cell. Compare the potential differences for the cells that you measured in your experiment with those that you have calculated and comment upon the accuracy and any differences that you observe.

Oxidation/Reduction equation worksheet

For the following six reactions, identify the oxidation numbers of the elements in the elements and compounds involved. Showing how you have calculated change in oxidation numbers of the elements. From this, **determine** which equations are examples of redox reactions and which are not.

1. $NaCl_{(aq)} + AgNO_{3\ (aq)} \rightarrow AgCl_{(s)} + NaNO_{3\ (aq)}$

2. $H_{2\ (aq)} + F_{2\ (aq)} \rightarrow 2HF_{(aq)}$

3. $PCl_{5\ (aq)} + 4H_2O_{(l)} \rightarrow H_3PO_{4\ (aq)} + 5HCl_{(aq)}$

4. $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\ (aq)} + 2H_{2\ (g)}$

5. $Ca(OH)_{2\ (aq)} + 2HCl_{(aq)} \rightarrow CaCl_{2\ (aq)} + 2H_2O_{(l)}$

6. $3CuS_{(s)} + 8HNO_{3\ (aq)} \rightarrow 3CuSO_{4\ (aq)} + 8NO_{(g)} + 4H_2O_{(l)}$

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